HW: WS on Bohr's Hydrogen Atom

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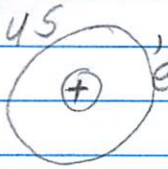
Chapter 4:

Electrons

Rutherford's Model - electron orbits around the nucleus, like planets rotate around the sun.



1908:

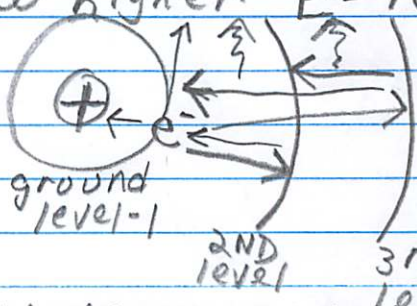
Rutherford is wrong because he has no explanation to why e⁻s do not collapse into the positive nucleus; and he does not explain chemical reactivity; why chemistry happens. Therefore, back to the drawing board.

1913, Neils Bohr - Model of H atom used the Bright Line Spectrum of Hydrogen to derive his model.

Each element has its own, unique Bright Line Spectrum, used by scientists to identify elements.

Bohr's Model - e⁻ circle the nucleus in 2D orbits, infinite # of orbitals

- When energy is added to the atom, the e⁻ absorbs the E and jumps to a higher E-level (excited State) e⁻ unstable at higher levels
- When e⁻ falls back down, gives off E in the form of light



the model only worked for Hydron →

Other element's Bright Line Spectrum Model did not match Niels Bohris Model and did not explain reactivity.

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Electrons and X-rays gave the same diffraction patterns. Electrons act like wave matter acts like Energy

Wave Particle Duality:
Electrons act like both waves and particles.

Schrödinger- had an equation that explains wave behavior; will use this equation to model e⁻ behavior.

Heisenberg- Uncertainty Principle:
1) we cannot know where an electron is and how fast it is traveling with any precision because the act of trying to find an e⁻ will change its behavior.

Quantum Model of the atom:
Gives the space with the best chance of finding where electrons are.

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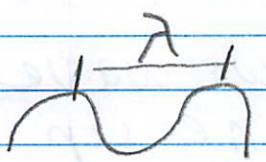
Spectrum Lines identify elements
Elements have their own unique spectrum

1923 - Louie de Broglia, X-ray defraction 1st to talk about electrons act weird. He suggested that electrons act like waves. Had calculations to back up this idea.

$E = mc^2$

$E = \frac{hc}{\lambda}$ ← speed of light

Plank's Constant



$$mc^2 = \frac{hc}{\lambda}$$

$$mc = \frac{h}{\lambda}$$

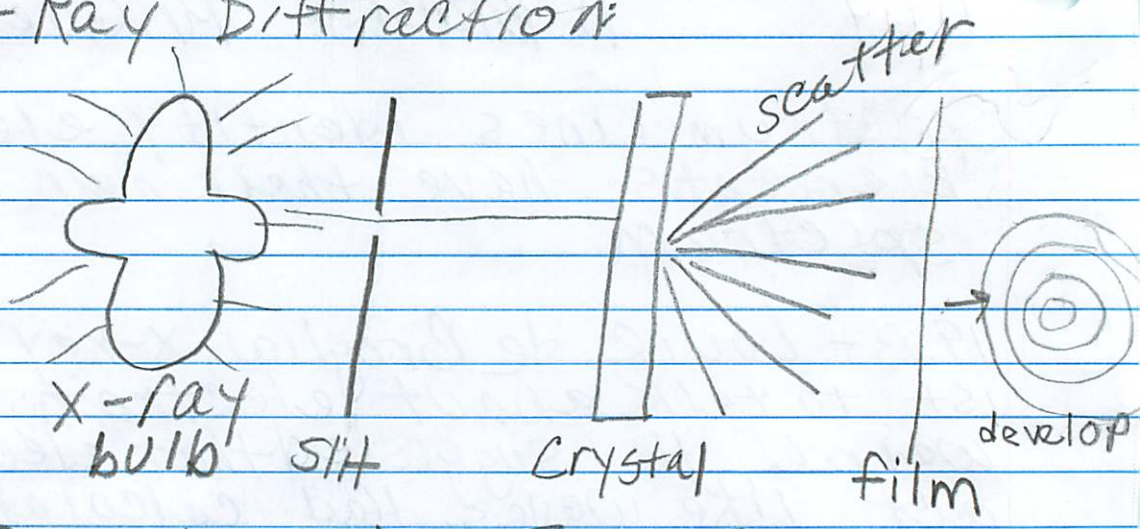
$$m = \frac{h}{\lambda c}$$

$$\lambda = \frac{h}{mc}$$

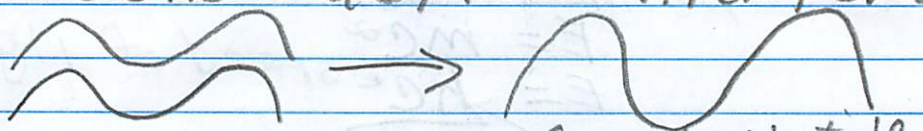
↑ speed

In theory, electrons have a wavelength same as X-rays. Proven by X-ray diffraction

X-Ray Diffraction

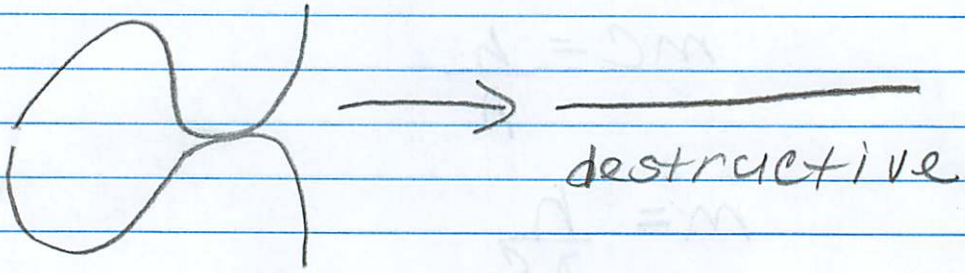


Constructive Interference

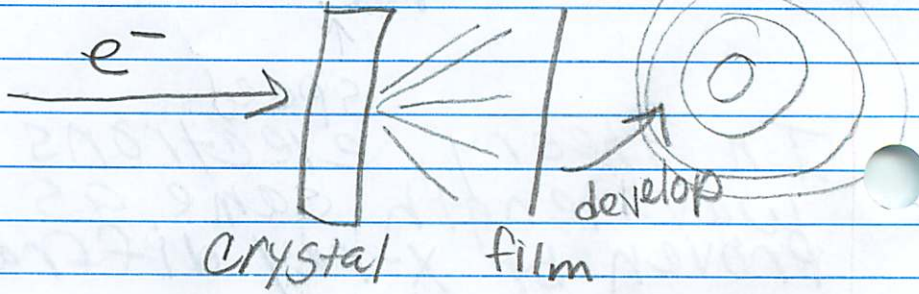


two waves line up → create bigger wave

Destructive Interference



Electron Beam ⇒ Pattern similar to X-rays



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Quantum Model of Atom vs. Bohr Model Compared and Contrasted

Bohr Model	Quantum Model
• e^- in 2D orbits; could pinpoint e^- location; • e^- where particles; • only works for Hydrogen; • did not explain chemical reactivity	• e^- in 3D orbitals which are the space with the highest chance of finding e^-; • e^- act like waves; • works for all elements; • explains chemical reactivity;

Compare: In both models, electrons can absorb energy, jump to higher energy levels and fall back down giving off light (creates a specific band of light).

Arrangement of Electrons According to Quantum Model:

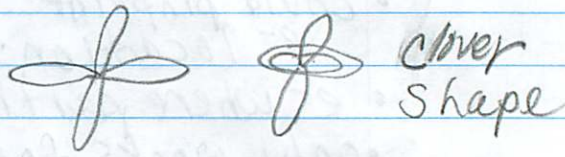
1. Principle Quantum # (n)
   ↳ tells you which energy level the electrons are in.
   $n = 1 \rightarrow \infty$ (7)

(over →)

2. Angular Momentum Quantum # (l)
↳ tells you which orbital the e^- are in.
l = s, p, d, f

3. Information about the orbitals:

orbital	Max # of e^-	Shape
s	2	sphere
p	6	propellar, or dumbbell
d	10	clover shape



4. Filling order of e^-:
Aufbau Principle:
electrons are added to Energy levels from low to high Energy

Low Energy 1 S

↑ Energy Level ↑ orbital

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Hw - Chapter homework due Fri. 11/18

Electron Arrangement in atoms:
e⁻ go from low → high energy.

- Principle Quantum number n is the energy level e⁻ is in.
  1st Energy level has only a "s" orbital.

- Orbitals ⇒ s, p, d, f

high Energy 4s 4p 4d
3s 3p 3d
2s 2p
Low Energy 1s ← orbital
↑ E level

Maximum # of e⁻ per orbital:
s = 2 e⁻
p = 6 e⁻
d = 10 e⁻

E level ↓ orbital

Filling Order - n ℓ

High E
5s 5p
4s 4p 4d
3s 3p 3d
2s 2p
1s
Low E ↑ ← orbital, ℓ
Energy Level, n

Hydrogen = 1 e⁻
e⁻ configuration:
$1s^1$

Lithium = 3 e⁻
$1s^2 2s^1$

Fluorine = 9 e⁻
$1s^2 2s^2 2p^5$

Silicon, Si, = 14 e^-

$1s^2 2s^2 2p^6 3s^2 3p^2$

Zinc, Zn, = 30 e^-

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$

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Valence e^- = outer shell e^-, #1 → 8
↳ highest E-level s & p e^-, 2 in s, 6 in p = 8
Core e^- = inner shell e^-

F: $1s^2 2s^2 2p^5$ - 7 valence e^-, 2 core e^-

Al: $1s^2 2s^2 2p^6 3s^2 3p^1$ - 3 valence e^-, 10 core

Na: $1s^2 2s^2 2p^6 3s^1$ - 1 valence e^-, 10 core e^-

Electron configuration Shorthand:
Noble Gas notation

	e^- configuration	Noble Gas notation
Na:	$\underline{1s^2 2s^2 2p^6} 3s^1$	[Ne] $3s^1$
Al:	$\underline{1s^2 2s^2 2p^6} 3s^2 3p^1$	[Ne] $3s^2 3p^1$
Cl:	$\underline{1s^2 2s^2 2p^6} 3s^2 3p^5$	[Ne] $3s^2 3p^5$

Neon's config.

Ca: [Ar] $4s^2$

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Valence e^-
↳ highest Energy level of s & p e^-

e^- dot conf.		e^- configuration	noble gas conf.	Val. e^-	Core e^-
·B·	B	$\underbrace{1s^2}_{\text{He conf.}}2s^2 2p^1$	$[He]2s^2 2p^1$	3	2
·Al·	Al	$\underbrace{1s^2 2s^2 2p^6}_{\text{Ne conf.}}3s^2 3p^1$	$[Ne]3s^2 3p^1$	3	10
K·	K	$\underbrace{1s^2 2s^2 2p^6 3s^2 3p^6}_{\text{Ar conf.}}4s^1$	$[Ar]4s^1$	1	18
·Fe·	Fe	$\underbrace{1s^2 2s^2 2p^6 3s^2 3p^6}_{\text{Ar conf.}}3d^6 4s^2$	$[Ar]3d^6 4s^2$	2	24